Chapter 2. Quantity of Heat

Presentation by

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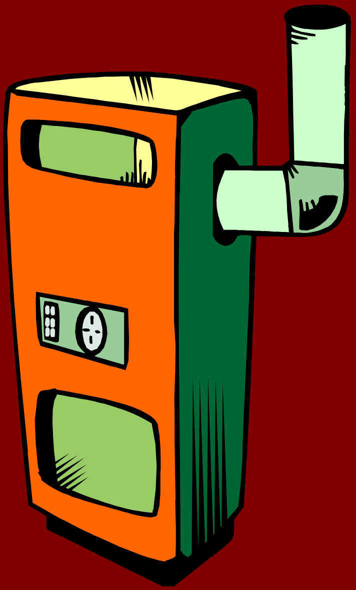
Objectives: After finishing this unit, you should be able to:

- Define the quantity of heat in terms of the calorie, the kilocalorie, the joule
- Write and apply formulas for specific heat capacity and solve for gains and losses of heat.
- Write and apply formulas for calculating the latent heats of fusion and vaporization of various materials.



Quantity of Heat

- The quantity of heat is the thermal energy required to raise the temperature of a given mass.
- Thermal energy is the energy associated with random molecular motion.
- Heat is defined as the transfer of energy across the boundary of a system due to a temperature difference between the system and its surroundings.
- We can measure changes in thermal energy by relating it to change in temperature.

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Heat and Internal Energy

- **At the outset, it is important that we make a major distinction between internal energy and heat.**
- **Internal energy** **is the all energy of a system that is associated with its microscopic components atoms and molecules.**

Units of Heat

One <u>calorie</u> (1 cal) is the quantity of heat required to raise the temperature of 1 g of water by 1 C^0.

<u>Example</u>

10 calories of heat will raise the temperature of 10 g of water by 10 C^0.

Units of Heat (Cont.)

One kilocalorie (1 kcal) is the quantity of heat required to raise the temperature of 1 kg of water by 1 C^0.

Example

10 kilocalories of heat will raise the temperature of 10 kg of water by 10 C^0.

The SI Unit of Heat

Since heat is energy, the joule is the preferred unit. Then, mechanical energy and heat are measured in the same fundamental unit.

Comparisons of Heat Units:

1 cal = 4.186 J

1 Btu = 778 ft lb

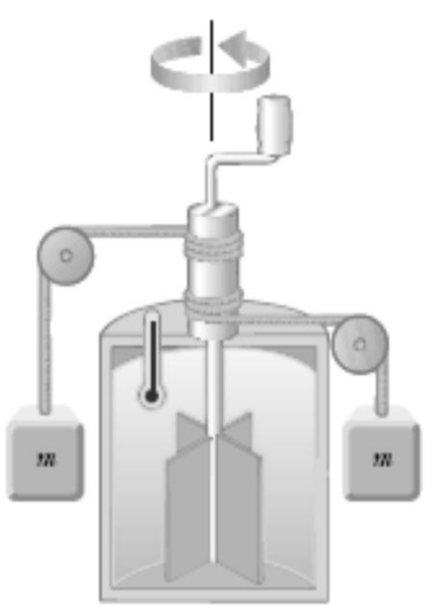
1 kcal = 4186 J

1 Btu = 252 cal

1 Btu = 1055 J

• ***Heat and Mechanical energy(Joule Experiment):***

- *Joule carried an experiment shown in the figure to give the relation between thermal energy units and mechanical energy units as shown:*



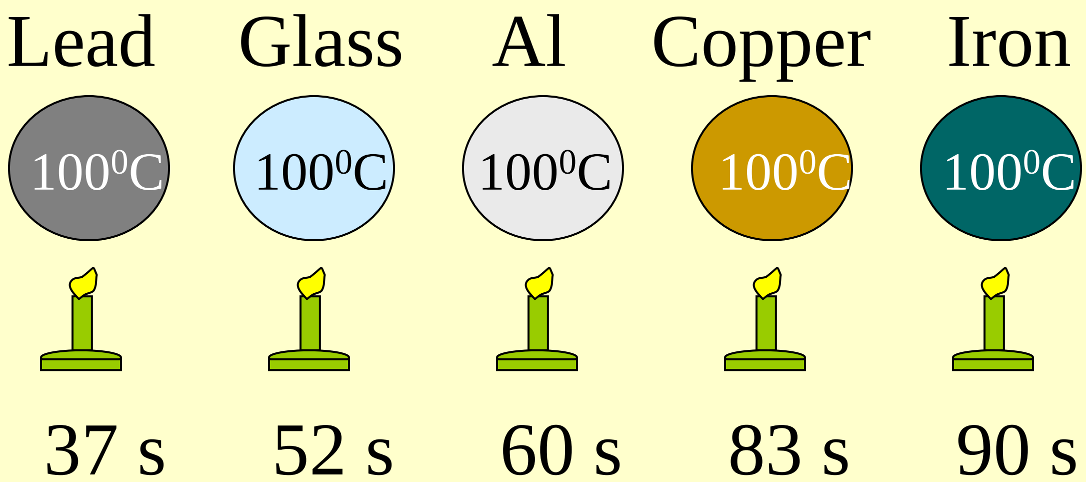
❑ **As the two blocks of mass m fall through a distance h, the loss in the potential energy (which equal the work done by paddles) is given by :**

- $W \propto Q \qquad , \qquad W = J \times Q$

- ***J is the mechanical heat equivalent defined as the work done to produce a quantity of heat equal to 1 cal.***

Heat Capacity

The heat capacity of a substance is the heat required to raise the temperature a unit degree.



Heat capacities based on time to heat from zero to 100^0C. Which has the greatest heat capacity?

Heact Capacity

Is the quantity of heat required to raise the temperature of the body one degree

$$\text{Heat Capacity} \quad S = \frac{Q}{\Delta T} \Leftarrow \frac{\text{Quantity of Heat}}{\text{Difference in Body Temperature}}$$

$$Units \Rightarrow \text{Cal}/{}^{0}\text{C},\ \text{kCal}/{}^{0}\text{C},\ \text{Btu}/{}^{0}\text{F},\ \text{J/K}$$

$$1\,\text{Btu} = 252\,\text{Cal},\ 1\ \text{J} \approx 0.239\,\text{Cal},\quad \text{and } 1\ \text{Cal} = 4.186\,\text{J}$$

Specific Heat Capacity

The specific heat capacity of a material is the quantity of heat needed to raise the temperature of 1 gm of object one degree.

$$c = \frac{Q}{m\Delta t}; \qquad Q = mc\Delta t$$

Water: $c = 1.0$ cal/g C^0 or 1 Btu/lb F^0 or 4186 J/kg K

Copper: $c = 0.094$ cal/g C^0 or 390 J/kg K

Comparison of Heat Units:

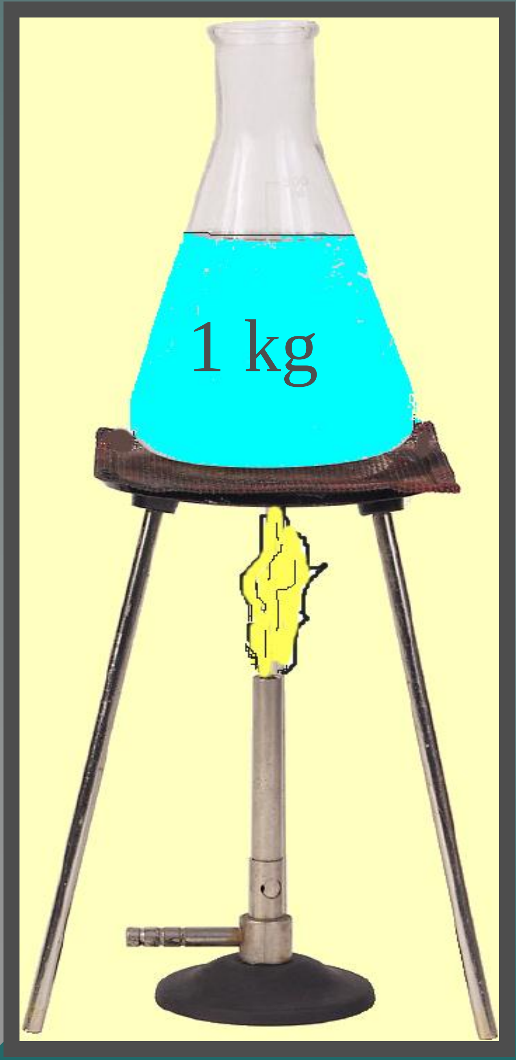
How much heat is needed to raise 1-kg of water from 0^0 to 100^0C?

The mass of one kg of water is: 1 kg = 1000 g = 0.454 lb_m

$$Q = mc\Delta t$$

1 lb_m = 454 g

For water: c = 1.0 cal/g C^0 or 1 Btu/lb F^0 or 4186 J/kg K



The heat required to do this job is:

10,000 cal   10 kcal

39.7 Btu   41, 860 J

Example

80 kg man ran a fever of 2 ^{0}C above normal; whose temperature was 39 ^{0}C inst ead of the normal 37 ^{0}C . Assuming that the human body is **<u>mostly water</u>**, how much heat is required to raise his temperature by amount?

Answer

$$\Delta T = 39^\circ C - 37^\circ C = 2^\circ C = 2\,K$$

$$\because \quad Q = m\,C\,\Delta T = (80\,kg)(4190\ J/kg.K)(2\,K)$$

$$= 6.7\,x10^5\,J$$

$$= 6.7\times10^5\times0.239\,cal = 160.13\ kCal$$

Problem

- How much heat is needed to raise the temperature of 30 kilograms of liquid water from 10 ^{0}C to 20 ^{0}C ?

- Soulution

- $Q = m\ c\ \Delta\ T =$ 30 Kg x 4.19 J/kg ^{0}C x (20-10)^{0}C
- = 1,257J

or

- $Q = m\ c\ \Delta\ Tc) =$ 30,000 x 1.0 cal/g ^{0}C x (20-10)^{0}C

= 30,000 gm x 1.0 cal/g ^{0}C x (20-10)^{0}C

= 300,000 cal

Problem Solving Procedure

1. Read problem carefully and draw a rough sketch.
2. Make a list of all given quantities
3. Determine what is to be found.
4. Recall applicable law or formula and constants.

$$c = \frac{Q}{m\Delta t}; \qquad Q = mc\Delta t$$

Water: $c = 1.0$ cal/g C^0 or 1 Btu/lb F^0 or 4186 J/kg K

5. Determine what was to be found.

Example 1: A 500-g copper coffee mug is filled with 200-g of coffee. How much heat was required to heat cup and coffee from 20 to 96^{0}C?

1. Draw sketch of problem.

2. List given information.

Mug mass $m_m = 0.500$ kg

Coffee mass $m_c = 0.200$ kg

Initial temperature of coffee and mug: $T_0 = 20^0$C

Final temperature of coffee and mug: $T_f = 96^0$C

3. List what is to be found:

Total heat to raise temperature of coffee (water) and mug to 96^{0}C.

Example 1(Cont.): How much heat needed to heat cup and coffee from 20 to 96^0C?
M_m = 500 kg ; m_C = 0.2 kg.

4. Recall applicable formula or law:

Heat Gain or Loss: $Q = mc\,\Delta t$

5. Decide that TOTAL heat is that required to raise temperature of mug and water (coffee). Write equation.

$$Q_T = m_m c_m \Delta t + m_c c_c \Delta t$$

Example 1(Cont.):

6. Substitute info and solve problem:

Copper: $c_m = 390$ J/kg C^0

Coffee (water): $c_c = 4186$ J/kg C^0

$$Q_T = m_m c_m \Delta t + m_c c_c \Delta t$$

Water: (0.20 kg)(4186 J/kgC^0)(76 C^0)

$\Delta t = 96^0C - 20^0C = 76\ C^0$

Cup: (0.50 kg)(390 J/kgC^0)(76 C^0)

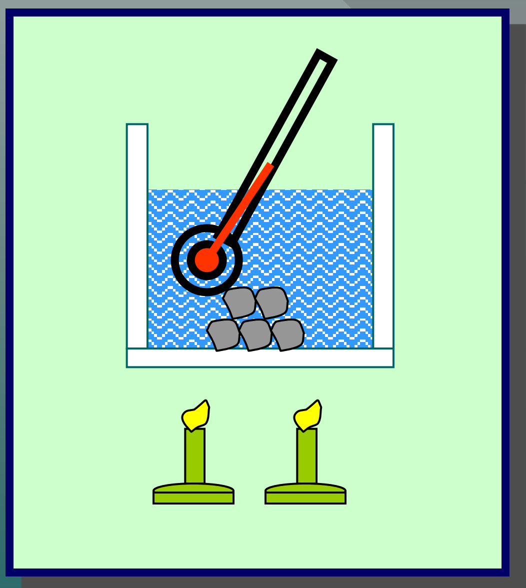
$Q_T = 63{,}600\ \text{J} + 14{,}800\ \text{J}$

$Q_T = 78.4$ kJ

A Word About Units

The substituted units must be consistent with those of the chosen value of specific heat capacity.

For example: **Water c_w = 4186 J/kg C^0 or 1 cal/g C^0**



$$Q = m_w c_w \Delta t$$

If you use 1 cal/g C^0 for *c*, then *Q* must be in calories, and *m* must be in grams.

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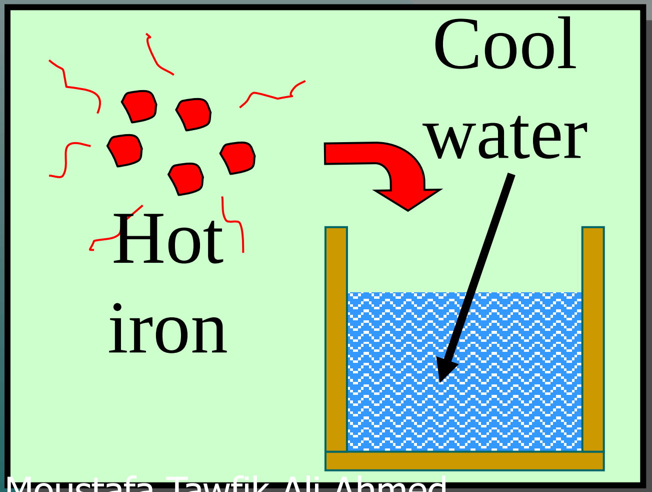
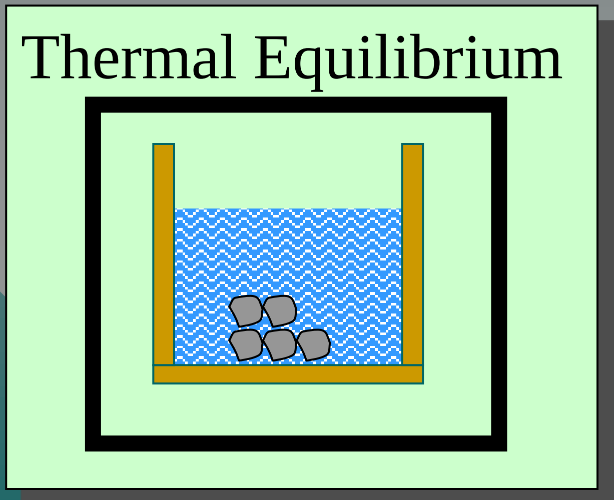
Conservation of Energy

Whenever there is a transfer of heat within a system, the heat lost by the warmer bodies must equal the heat gained by the cooler bodies:

$$\Sigma \text{ (Heat Losses)} = \Sigma \text{ (Heat Gained)}$$





Example: ***A handful of copper shot heated to 90°C and then dropped into* 80 g *of water at 10 °C. The final temperture of the mixture is 18°C. What was the mass of the shot? Note: C (shot)=0.093*** **cal/g °C**

Answer

Copper shot temperature $(T_s) = 90^{o}C$,

water temperature $(T_w) = 10^{o}C$,

Equilibrium temperature (Te) $= 18^{0}C$,

$m_w = 80g, \quad m_s = ???$

Total Heat Lost (by shot) = Total Heat Gained (by water)

$$Q_L = Q_G \Rightarrow m_s C_s \Delta T_s = m_w C_w \Delta T_w$$

$$\Delta T_s = T_s - T_e = 90\,^{0}C - 18\,^{0}C = 72\,^{0}C$$

$$\Delta T_w = T_e - T_w = 18\,^{0}C - 10\,^{0}C = 8\,^{0}C$$

m_s (0.093) (72) = (80) (1) (8) ⟶

$$\therefore \quad \boldsymbol{m_s = 95.6\,g}$$

Example

Eighty grams of dry iron shot is placed in a cup and heated to a temperature of 95°C. The mass of the inner aluminum cup and the aluminum stirrer is 60g. The calorimeter is partially filled with 150g of water at 18°C. The hot shot is quickly poured into the cup, and the calorimeter is sealed, after the system has reached thermal equilibrium, the final temperature is 22°C . Compute the specific heat of the iron.

Answer

Total Heat Lost (by Iron) = Total Heat Gained (by water+Aluminum Container)

$$Q_L\,(\text{by Iron}) = Q_G\,(\text{by water} + \text{Alum.})$$

Iron shot temp. $(T_I) = 95°C$, water temp. $(T_w) = 18°C$, Alum. temp. $(T_{Al}) = 18°C$,

Equilibrium temp. (Te) = 22^0C, $m_I = 80g$, $m_w = 150g$, $m_{Al} = 60g$, $C_{Al} = 0.22\,Cal/g.^0C$

$$C_I = ????\,cal/g.^0C$$

$$\Delta T_I = T_I - T_e = 95\,^0C - 22\,^0C = 73\,^0C$$

$$\Delta T_w = \Delta T_{Al} = T_e - T_w = 22\,^0C - 18\,^0C = 4\,^0C$$

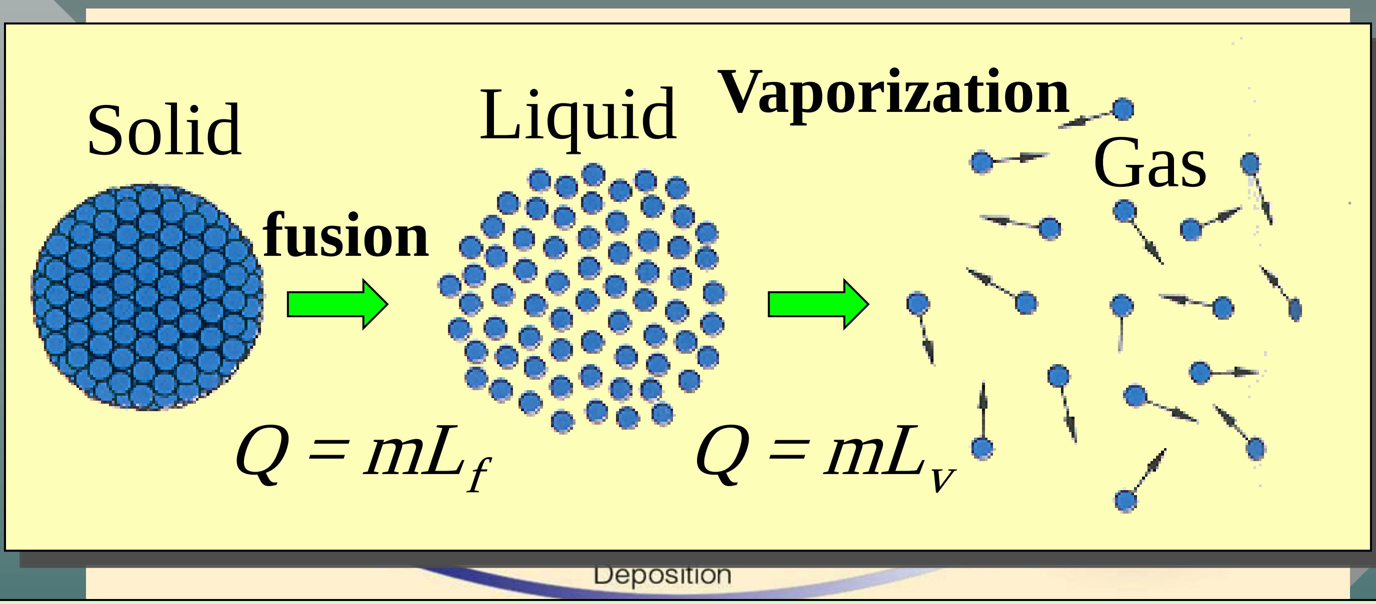
$$m_I C_I \Delta T_I = (m_w C_w + m_{Al} C_{Al})\Delta T_w$$

$$80\,C_I\,(73) = (150 \times 1 + 60 \times 0.22)(4)$$

$$C_I = 0.1117\ Cal / g.^0C$$

Change of Phase

When a change of phase occurs, there is only a change in potential energy of the molecules. The temperature is constant during the change.



Terms: , latent heats, Fusion, vaporization, condensation, Sublimation, evaporation, freezing point, melting point.

Change of Phase

Generally: The amount of heat needed per unit mass to produce a phase change is called the *latent heat (L)*

The latent heat of fusion (L_f) The amount of heat needed to change one gram of a <u>solid</u> substance into one gram of <u>liquid</u> without changing its temperature

$$L_f = \frac{Q}{m}$$

For Water: L_f = 80 cal/g = 333,000 J/kg

The latent heat of vaporization (L_v) The amount of heat needed to vaporize a unit mass of liquid

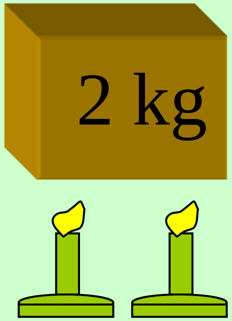
$$L_v = \frac{Q}{m}$$

For Water: L_v = 540 cal/g = 2,256,000 J/kg

Melting a Cube of Copper

The heat Q required to melt a substance at its melting temperature can be found if the mass and latent heat of fusion are known.

$$Q = mL_v$$



What Q to melt copper?

L_f = 134 kJ/kg

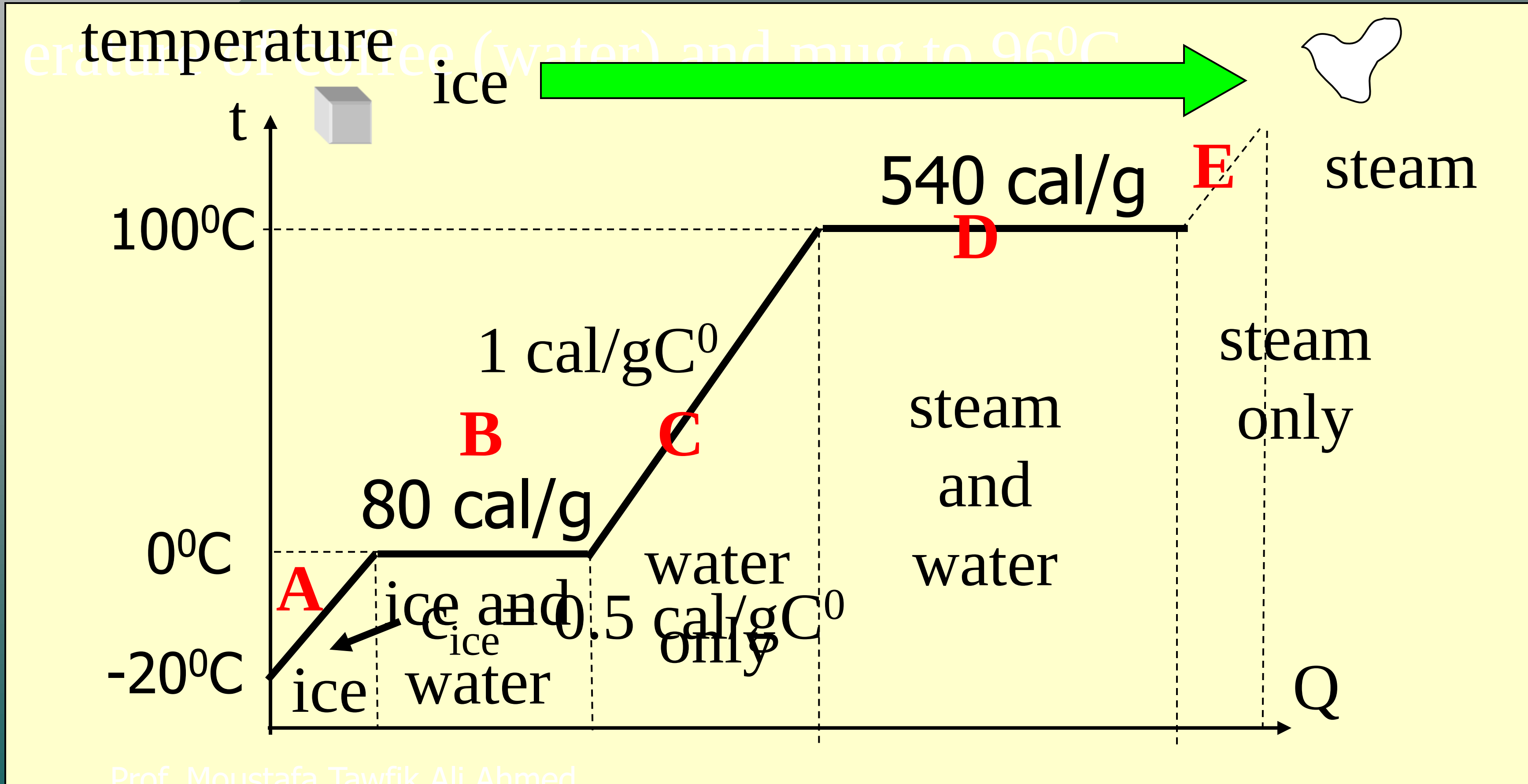
<u>Example:</u> To completely melt 2 kg of copper at 1040^{0}C, we need:

$$Q = mL_f = (2\text{ kg})(134{,}000\text{ J/kg})$$

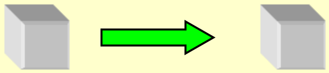
$$Q = 268\text{ kJ}$$

Example 3: How much heat is needed to convert 10 g of ice at -20^{0}C to steam at 100^{0}C?

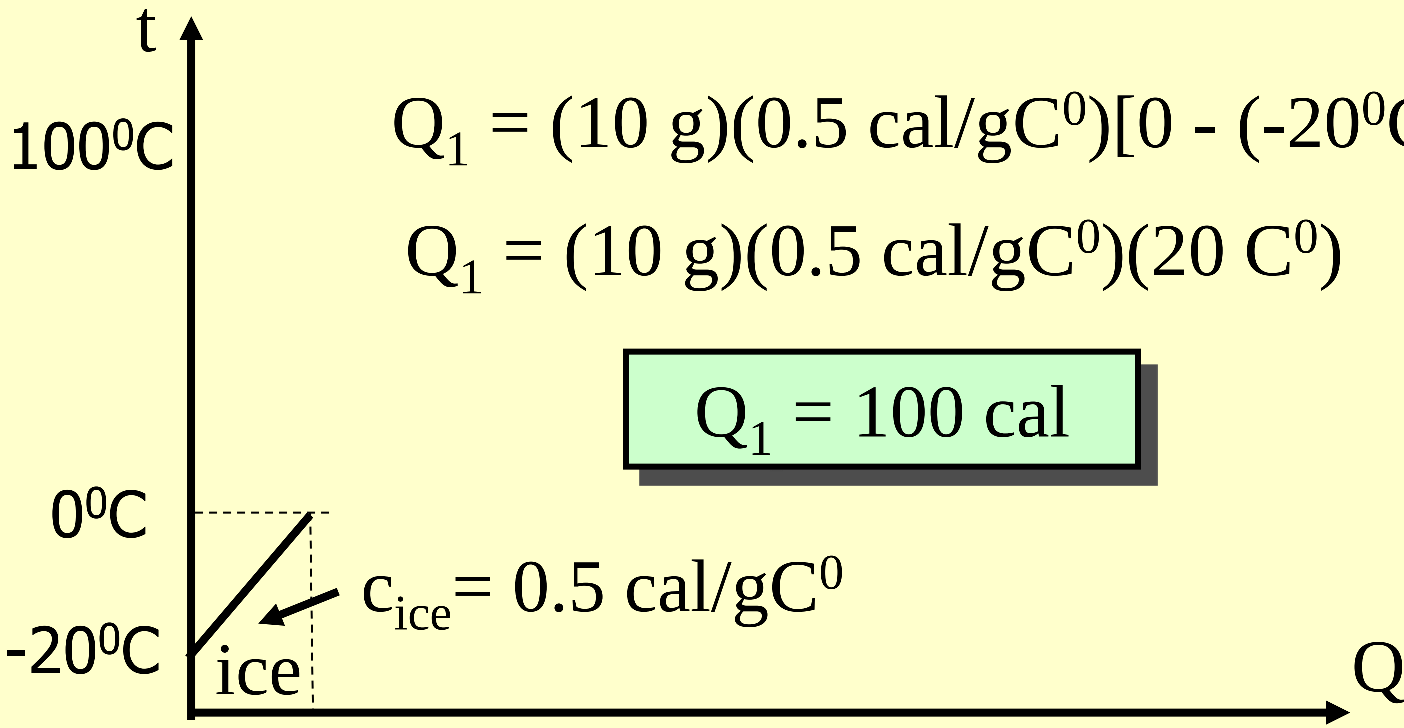
First, let's review the process graphically as shown:



Example 3 (Cont.): Step one is Q_1 to convert 10 g of ice at -20^0C to ice at 0^0C (no water yet).



Q_1 to raise ice to 0^0C: $Q_1 = mc\Delta t$

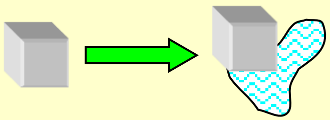


$$Q_1 = (10\ g)(0.5\ cal/gC^0)[0 - (-20^0C)]$$

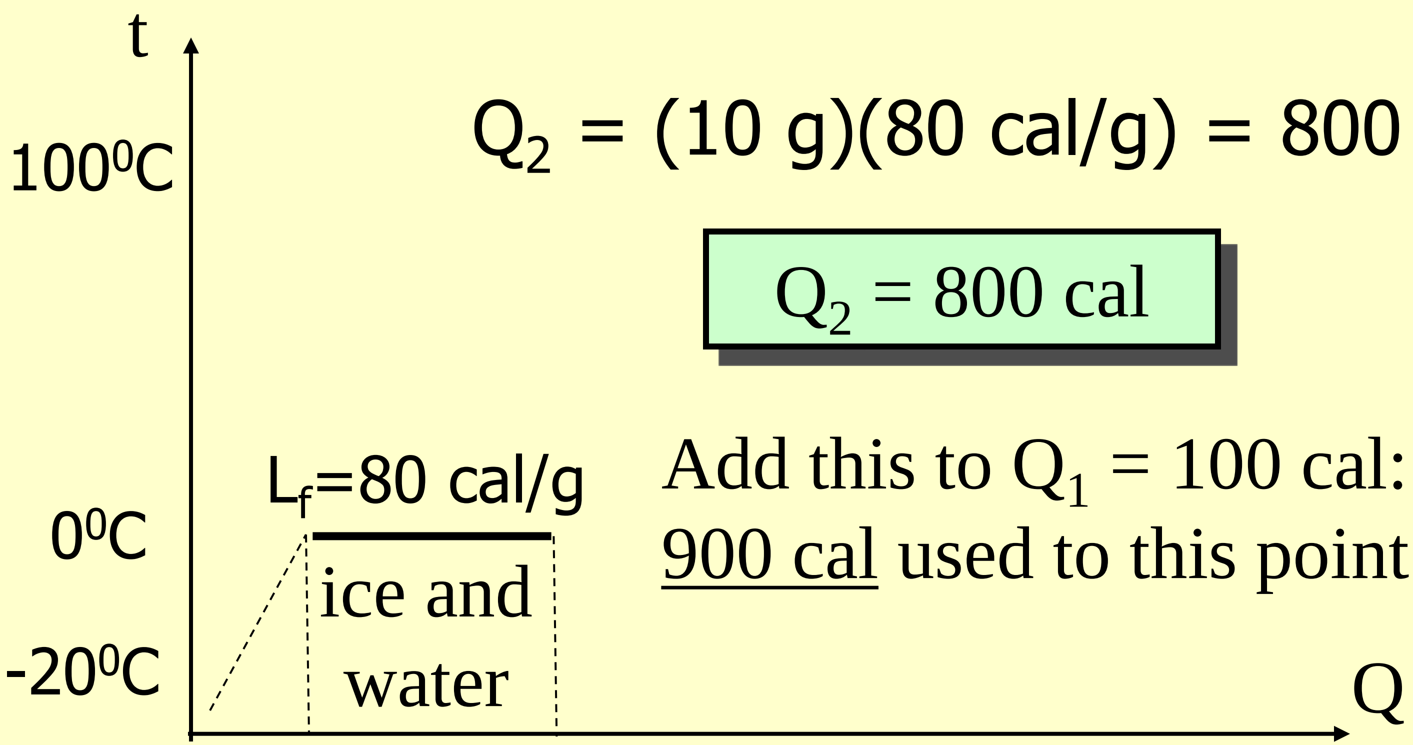
$$Q_1 = (10\ g)(0.5\ cal/gC^0)(20\ C^0)$$

$$\boxed{Q_1 = 100\ cal}$$

Example 3 (Cont.): Step two is Q_2 to convert 10 g of ice at 0^0C to water at 0^0C.



Q_2 to melt 10 g of ice at 0^0C: $Q_2 = mL_f$

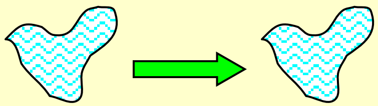


$$Q_2 = (10\text{ g})(80\text{ cal/g}) = 800\text{ cal}$$

$$\boxed{Q_2 = 800\text{ cal}}$$

Add this to Q_1 = 100 cal: <u>900 cal</u> used to this point.

Example 3 (Cont.): Step three is Q_3 to change 10 g of water at 0^0C to water at 100^0C.

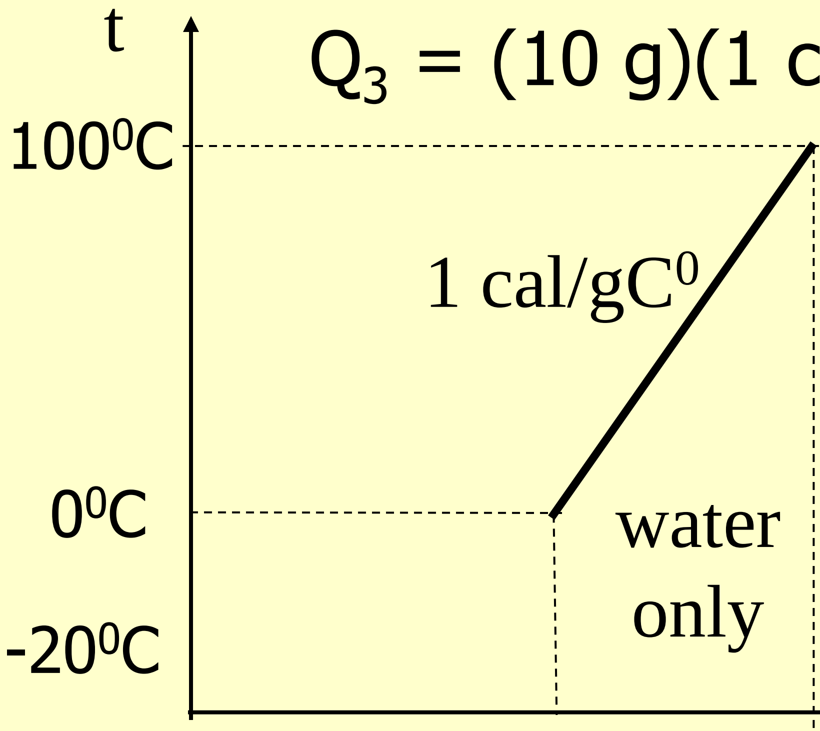


0^0C to 100^0C

Q_3 to raise water at 0^0C to 100^0C.

$Q_3 = mc\Delta t\,; \quad c_w = 1\ \text{cal/gC}^0$

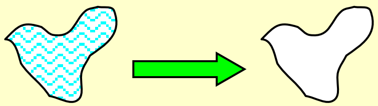
$$Q_3 = (10\ \text{g})(1\ \text{cal/gC}^0)(100^0\text{C} - 0^0\text{C})$$



$$\boxed{Q_3 = 1000\ \text{cal}}$$

$$\text{Total} = Q_1 + Q_2 + Q_3$$
$$= 100 + 900 + 1000$$
$$= 1900\ \text{cal}$$

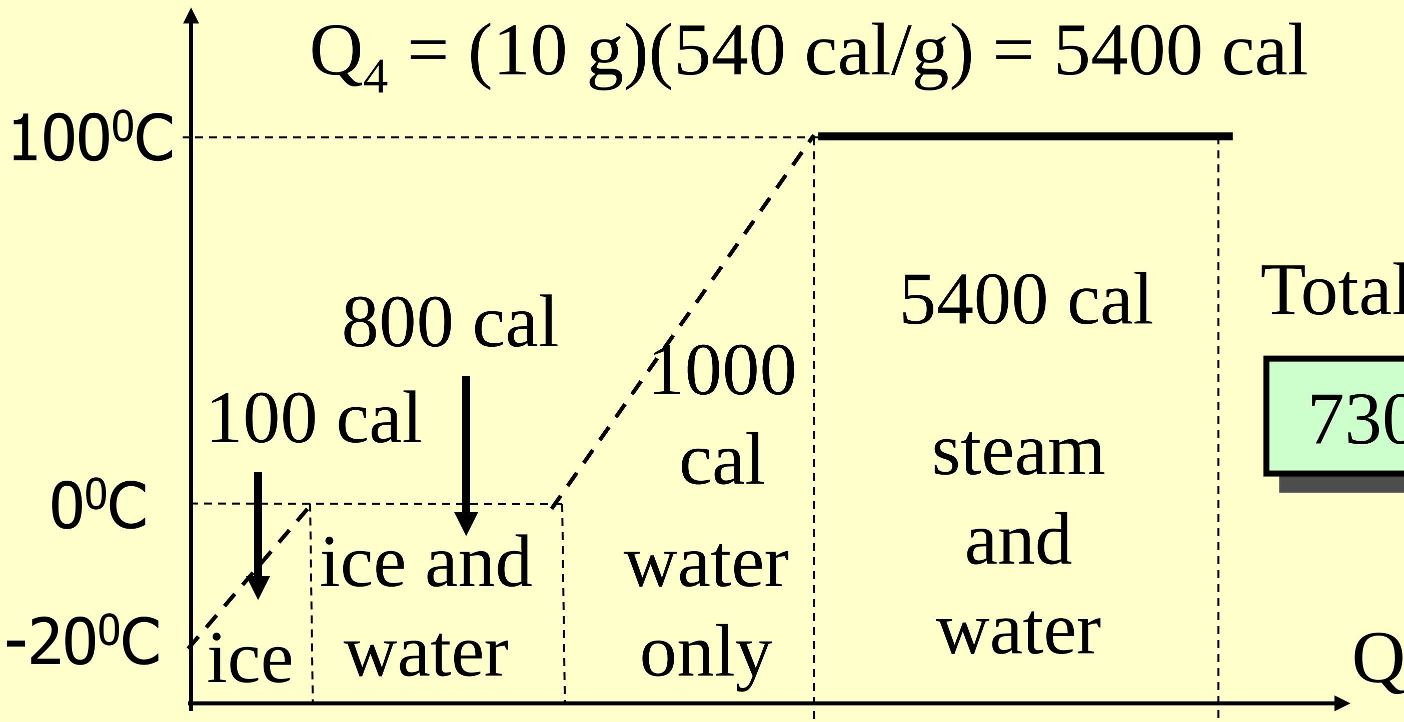
Example 3 (Cont.): Step four is Q_4 to convert 10 g of water to steam at 100^0C? ($Q_4 = mL_v$)



vaporization

Q_4 to convert all water at 100^0C to steam at 100^0C. ($Q = mL_v$)

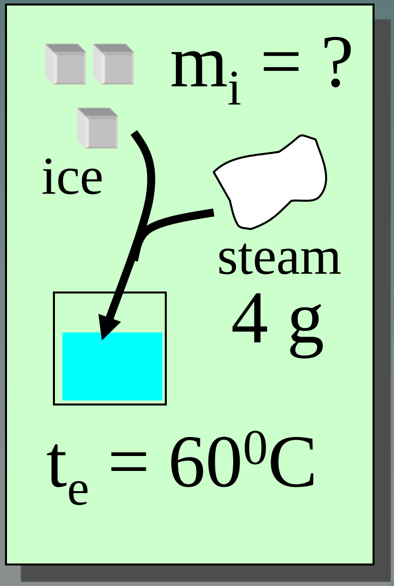
$$Q_4 = (10\text{ g})(540\text{ cal/g}) = 5400\text{ cal}$$



Total Heat:

7300 cal

Example 4: How many grams of ice at 0^0C must be mixed with 4 gm of steam in order to produce water at 60^0C?



Ice must melt and then rise to 60^0C.

Steam must condense and drop to 60^0C.

Total Heat Gained = Total Heat Lost

$$m_iL_f + m_ic_w\Delta t = m_sL_v + m_sc_w\Delta t$$

Note: All losses and gains are absolute values (positive).

Total Gained: $m_i(80\ cal/g) + m_i(1\ cal/gC^0)(60\ C^0 - 0^0C)$

Total Lost: $(4\ g)(540\ cal/g) + (4\ g)(1\ cal/gC^0)(100^0C - 60^0C)$

Example 4 (Continued)

Total Gained: $= 80m_i + 60m_i$

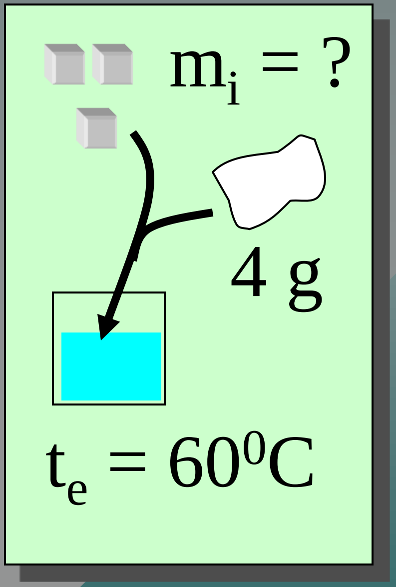
Total Lost: $= 2160 \text{ g} + 160 \text{ g}$

Total Heat Gained = Total Heat Lost

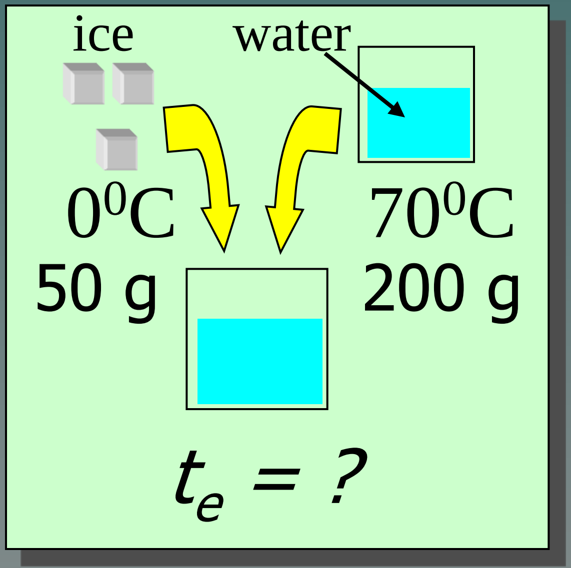
$$80m_i + 60m_i = 2160 \text{ g} + 160 \text{ g}$$

$$m_i = \frac{2320 \text{ g}}{140}$$

$$\boxed{m_i = 16.6 \text{ g}}$$



Example 5: Fifty grams of ice are mixed with 200 g of water initially at 70^0C. Find the equilibrium temperature of the mixture.



Ice melts and rises to t_e

Water drops from 70 to t_e.

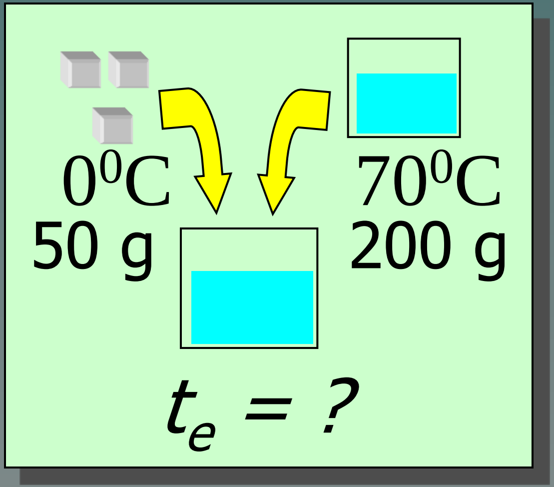
Heat Gained: $m_i L_f + m_i c_w \Delta t$; $\Delta t = t_e - 0^0$C

Gain = (50 g)(80 cal/g) + (50 g)(1 cal/gC^0)(t_e - 0^0C)

Gain = 4000 cal + (50 cal/g)t_e

Example 5 (Cont.):

$$\text{Gain} = 4000 \text{ cal} + (50 \text{ cal/g})t_e$$



$$\text{Heat Lost} = m_w c_w \Delta t$$

$$\Delta t = 70^0\text{C} - t_e \quad \text{[high - low]}$$

$$\text{Lost} = (200 \text{ g})(1 \text{ cal/gC}^0)(70^0\text{C} - t_e)$$

$$\text{Lost} = 14{,}000 \text{ cal} - (200 \text{ cal/C}^0)\ t_e$$

Heat Gained Must Equal the Heat Lost:

$$4000 \text{ cal} + (50 \text{ cal/g})t_e = 14{,}000 \text{ cal} - (200 \text{ cal/C}^0)\ t_e$$

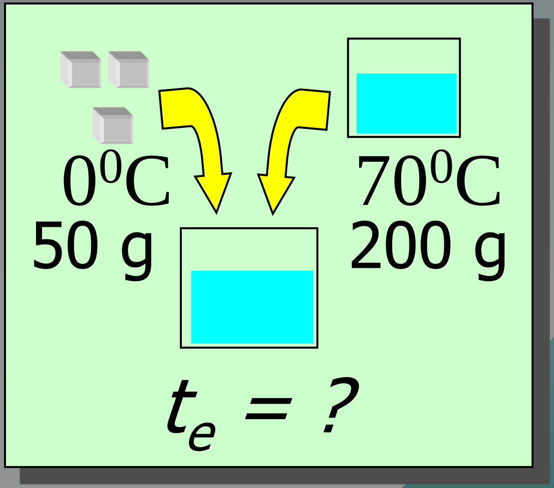
Example 5 (Cont.):

Heat Gained Must Equal the Heat Lost:

4000 cal + (50 cal/g)t_e = 14,000 cal - (200 cal/C^0) t_e

Simplifying, we have: (250 cal/C^0) t_e = 10,000 cal

$$t_e = \frac{10,000\text{ cal}}{250\text{ cal/C}^0} = 40^0\text{C}$$



$$\boxed{t_e = 40^0\text{C}}$$

Summary of Heat Units

One calorie (1 cal) is the quantity of heat required to raise the temperature of 1 g of water by 1 C^0.

One kilocalorie (1 kcal) is the quantity of heat required to raise the temperature of 1 kg of water by 1 C^0.

One British thermal unit (Btu) is the quantity of heat required to raise the temperature of 1 lb of water by 1 F^0.

Summary: Change of Phase

The **latent heat of fusion** (L_f) of a substance is the heat per unit mass required to change the substance from the solid to the liquid phase of its melting temperature.

$$L_f = \frac{Q}{m}$$

For Water: L_f = 80 cal/g = 333,000 J/kg

The **latent heat of vaporization** (L_v) of a substance is the heat per unit mass required to change the substance from a liquid to a vapor at its boiling temperature.

$$L_v = \frac{Q}{m}$$

For Water: L_v = 540 cal/g = 2,256,000 J/kg

Summary: Specific Heat Capacity

The specific heat capacity of a material is the quantity of heat to raise the temperature of a unit mass through a unit degree.

$$c = \frac{Q}{m\Delta t}; \qquad Q = mc\Delta t$$

Summary: Conservation of Energy

Whenever there is a transfer of heat within a system, the heat lost by the warmer bodies must equal the heat gained by the cooler bodies:

$$\Sigma \text{ (Heat Losses)} = \Sigma \text{ (Heat Gained)}$$

Summary of Formulas:

$$c = \frac{Q}{m\Delta t}; \qquad Q = mc\Delta t$$

$$\Sigma \text{ (Heat Losses)} = \Sigma \text{ (Heat Gained)}$$

$$L_f = \frac{Q}{m}; \qquad Q = mL_f$$

$$L_v = \frac{Q}{m}; \qquad Q = mL_v$$

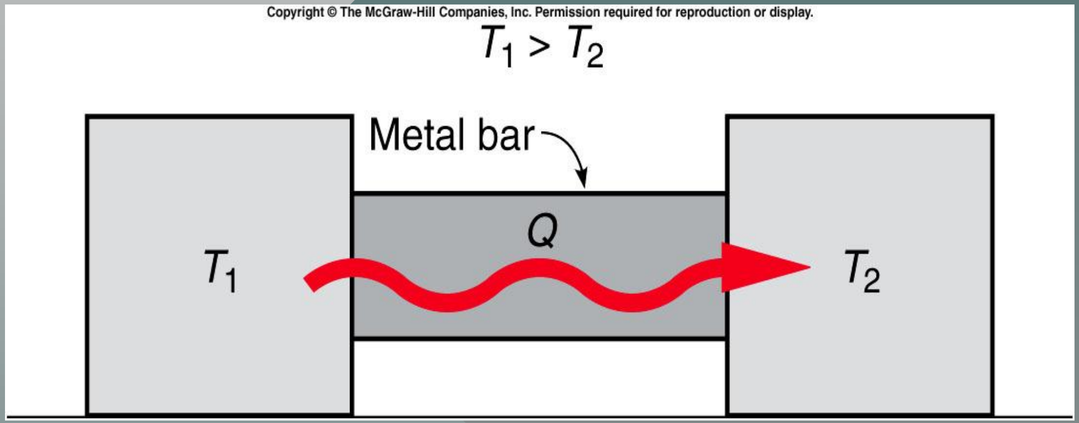
General knowledge

- Temperature is a measure of the activity of molecules.
- Heat is a measure of energy
- Specific Heat determines the amount of energy that a substance stores
- Heat of Fusion or Vaporization is a measure of the energy stored in the physical state of matter

The Flow of Heat

- Three basic processes for heat flow:
  - Conduction
  - Convection
  - Radiation

Conduction: heat flow when in contact



- **The rate of heat flow depends on:**

  a) the temperature difference between the objects.

  b) The *thermal conductivity* of the materials, a measure of how well the materials conduct heat.

– A metal block at room temperature will feel colder than a wood block of the exact same temperature. Why?

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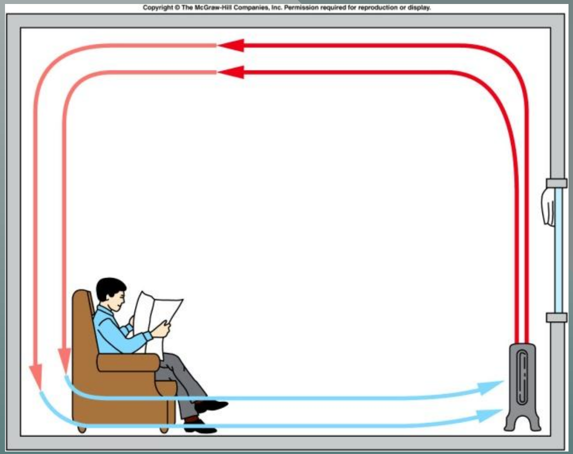
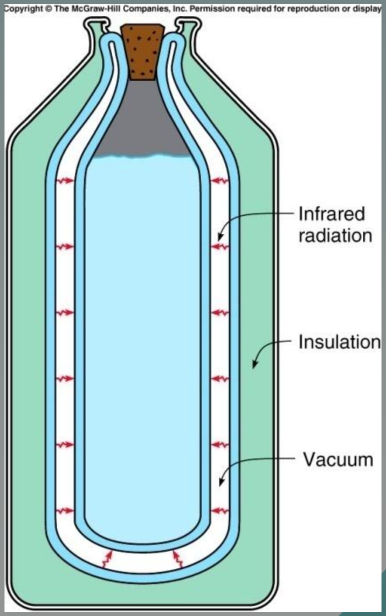
Convection: **heat is transferred by the motion of a fluid containing thermal energy.**



Radiation, **heat energy is transferred by electromagnetic waves.**

can take place across a vacuum.



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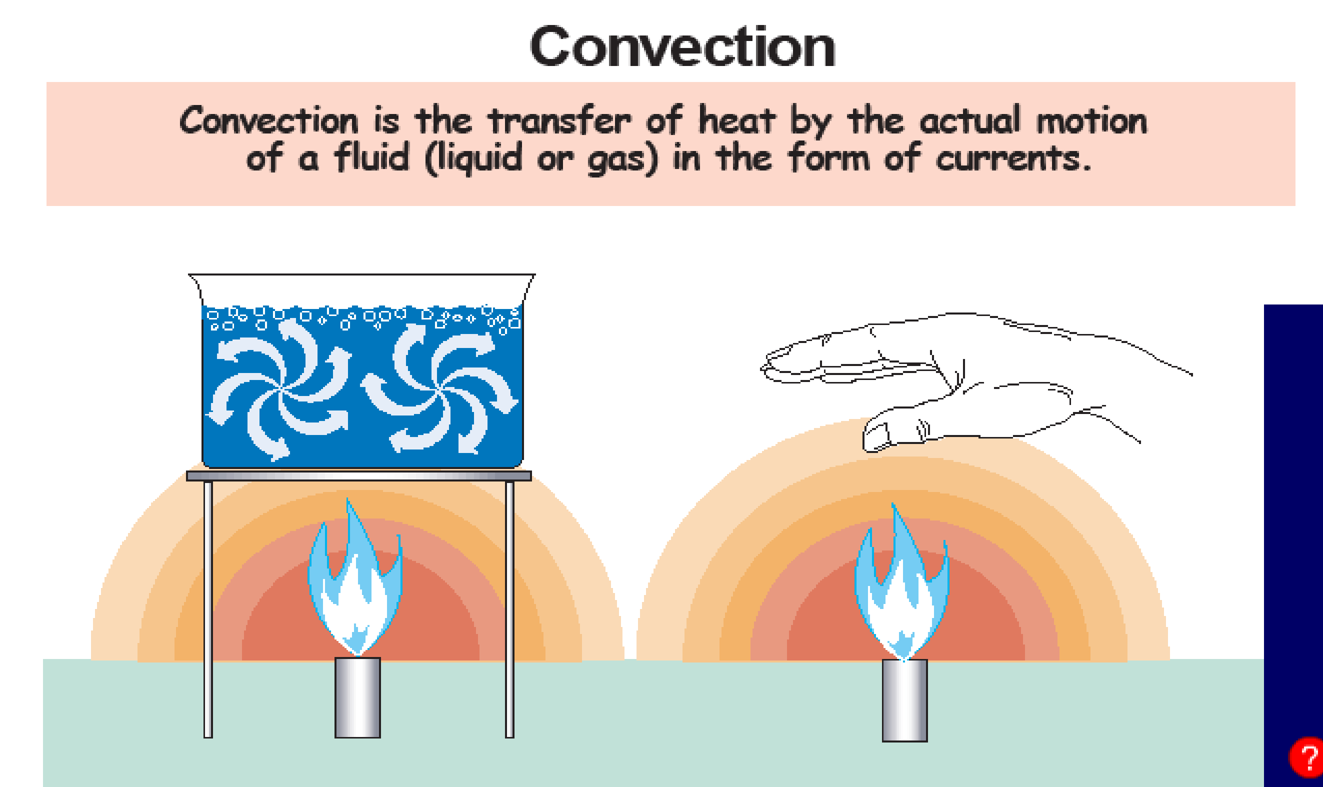

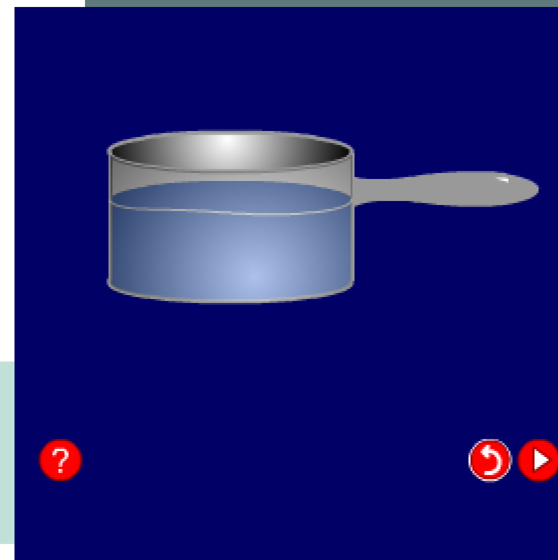

Convection

Due to difference in temperature, causing fluids to move in a circular motion

Example hot air rises, cool air sinks

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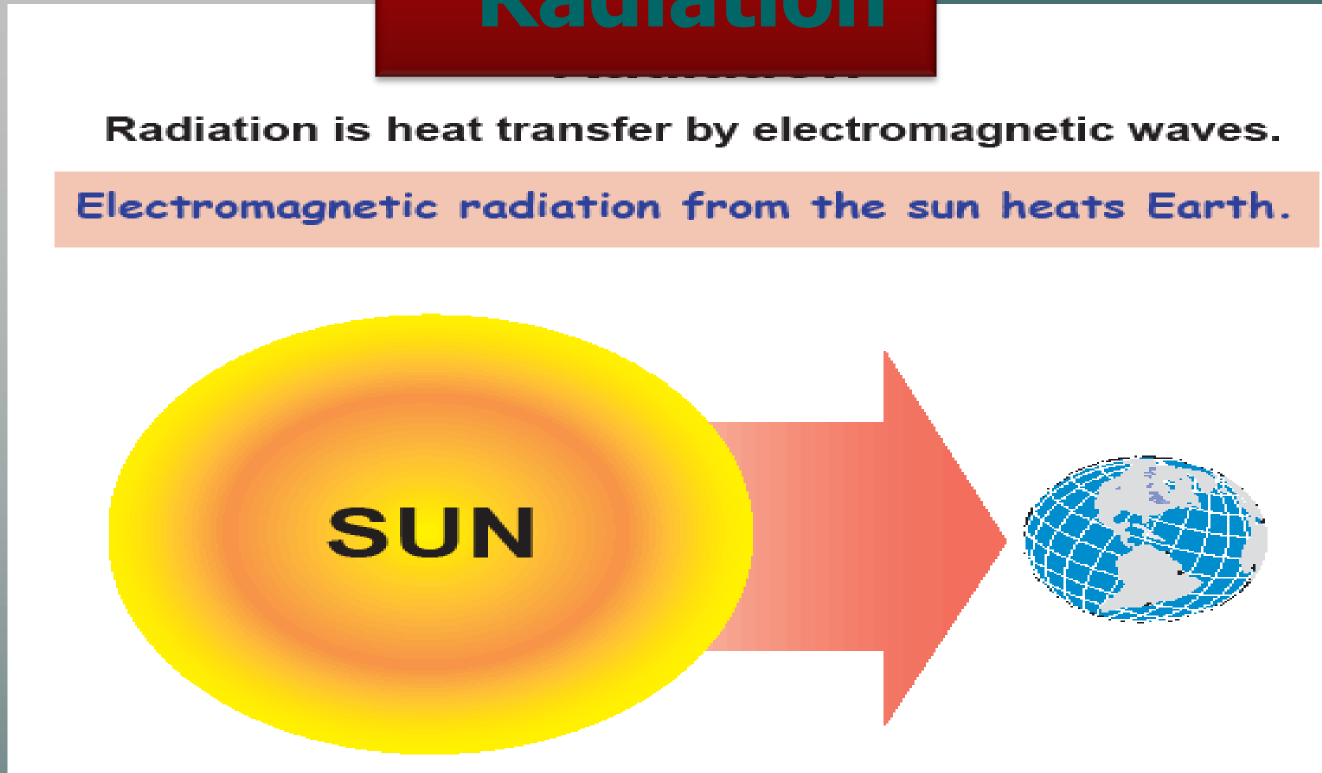
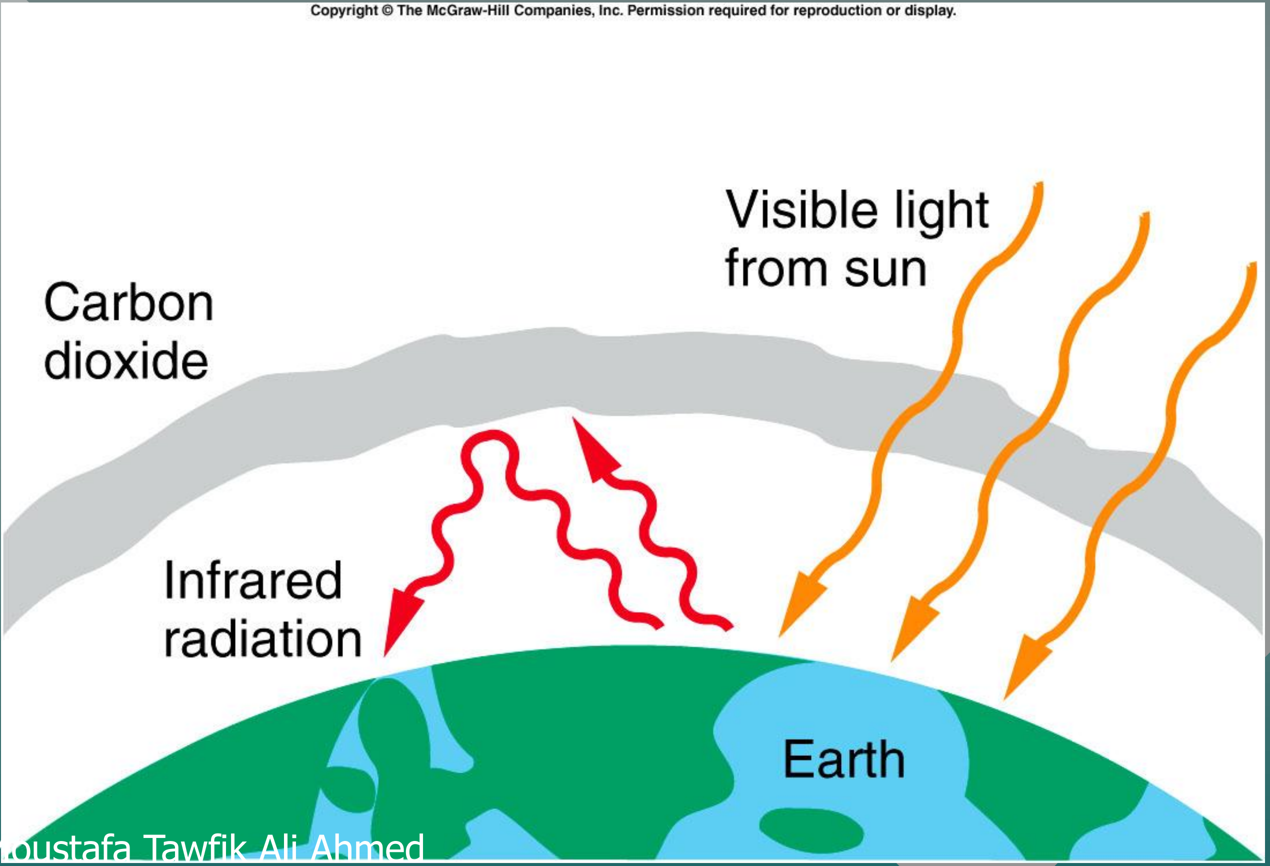
Radiation



- Radiant energy or heat

- Produced by the internal vibration of the particles of a body that is a source of heat. It then moves through space like a wave.

What heat-flow processes are involved in the greenhouse effect?



Some Concepts

Radiation travels in straight lines

True/False

Radiation can travel through a vacuum

True/False

Radiation requires particles to travel

True/False

1. Which of the following is not a method of heat transfer?

A. Radiation

B. Insulation

C. Conduction

D. Convection

2. In which of the following are the particles closest together?

A. Solid

B. Liquid

C. Gas

D. Fluid

3. How does heat energy reach the Earth from the Sun?

A. Radiation
B. Conduction
C. Convection
D. Insulation

4. Which is the best surface for reflecting heat radiation?

A. Shiny white
B. Dull white
C. Shiny black
D. Dull black

5. Which is the best surface for absorbing heat radiation?

A. Shiny white
B. Dull white
C. Shiny black
D. Dull black

Practice Problem

- When you step out of a swimming pool on a hot, dry day, you feel quite chilly. Why?

- The water will start evaporating in the air. Evaporation requires energy (heat) and that heat will come from your body. Since you are losing heat, you will feel cold.

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